TOPIC 09 – REDOX

9.1 – INTRO TO OXIDATION AND REDUCTION

IB Chemistry

T09D01

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9.1 – Intro to Redox

- 9.1.1 Define oxidation and reduction in terms of electron loss and gain. (1)
- 9.1.2 Deduce the oxidation number of an element in a compound. (3)
- 9.1.3 State the names of compounds using oxidation numbers. (1)
- 9.1.4 Deduce whether an element undergoes oxidation or reduction in reactions using oxidation numbers. (3)

Redox Definitions:

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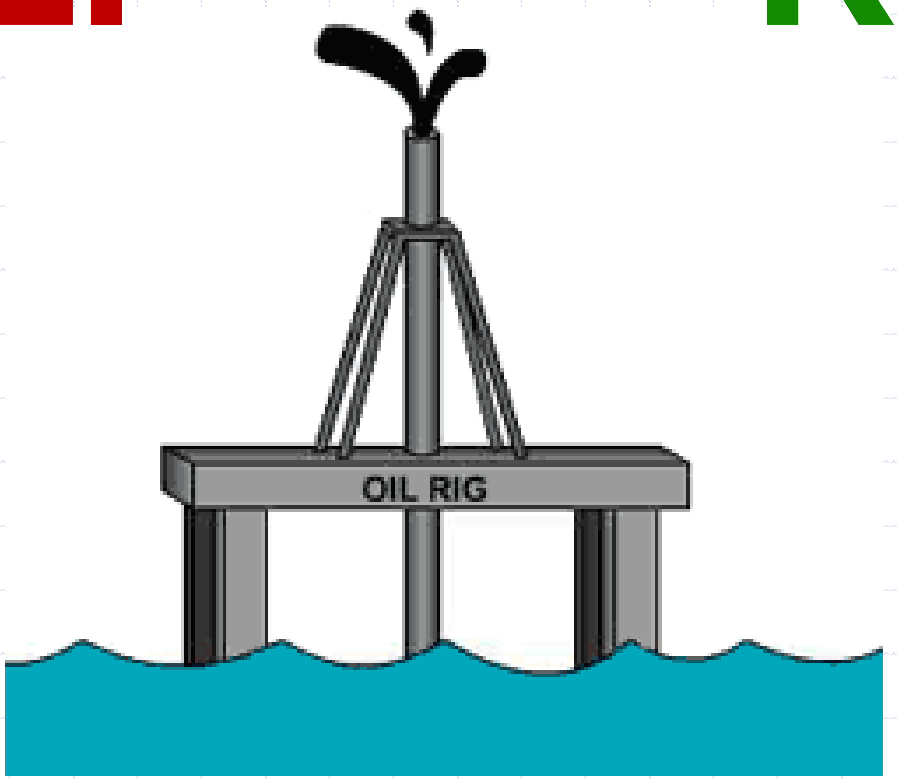
9.1.1 Define oxidation and reduction in terms of electron loss and gain. (1)

- Oxidation was *originally* defined as:
  - The addition of oxygen to a substance
    - $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
  - The loss or removal of hydrogen from a substance
    - $MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$
- Now, we consider **Oxidation** to be the loss of electrons for an element where it becomes more positive, or cationic.
  - $Mg \rightarrow Mg^{2+} + 2e^-$ (where Mg loses 2 electrons)
  - $Cl^- \rightarrow \frac{1}{2} Cl_2 + 1e^-$ (where Cl loses 1 electron)
- There must then be another part of the equation, the **Reduction**, or gain of electrons by an element where it becomes less positive, or anionic.

O.I.L. R.I.G.



OIL: Oxidation is Loss....
RIG: Reduction is Gain....
....of electrons

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Simple Half-Reactions

- Look at the same examples again:
- $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
  - Oxidation: $2x(Mg \rightarrow Mg^{2+} + 2e^-)$
  - Reduction: $O_2 + 4e^- \rightarrow 2O^{2-}$
- $MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$
  - Oxidation: $2x(Cl^- \rightarrow ½\ Cl^2 + 1e^-)$
  - Reduction: $Mn^{4+} + 2e^- \rightarrow Mn^{2+}$
- The same number of electrons must be gained by one (or more) species as is lost by another
- In order to understand and work with Redox Equations we must first be able to identify the oxidation state of each element

Oxidation Numbers:

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9.1.2 Deduce the oxidation number of an element in a compound. (3)

- The modern definition of redox reactions is very applicable to those including ionic compounds, such as:
  - $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
    - $2x\ (Na \rightarrow Na^+ + 1e^-)$
    - $2x\ (\frac{1}{2}\ Cl_2 + 1e^- \rightarrow Cl^-)$
- But what about equations not involving ionic species, such as:
  - $S(s) + O_2(g) \rightarrow SO_2(g)$
    - According to the original definition, S has been oxidized (gained O)
    - As you can see, no clear charge can be determined and electrons are not transferred from one species to another according to original convention.
    - The **oxidation number** must be assigned and monitored in order to make sense of this reaction

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Rules for Oxidation Numbers

- Ox # of an **uncombined element** is zero
  - O_2, Mg, Fe, H_2, Cl_2, Ag, etc
- For **simple ions**, the Ox # is the charge on that ion
  - Fe^{2+} (is +2), Cl^- (is +1), Ni^{3+} (is +3)
- For a **compound**, the sum of Ox #'s is zero
  - NaCl (Na^+ and Cl^- where +1 and -1 cancel)
- For an **oxoanion** (polyatomic ion) the sum is the charge
  - SO_4^{2-} [S is +6, 4x(O is -2)] = -2
- Ox # of **Hydrogen** is +1 (unless with a reactive metal)
- Ox # of **Oxygen** is -2
  - Exceptions: H_2O_2 (O is -1) and OF_2 (O is +2)

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Oxidation Numbers for Redox

- So, for our example of $S(s) + O_2(g) \rightarrow SO_2(g)$
  - S is alone, so it's zero
  - O_2 is alone, so it's zero
  - In SO_2, O is -2 and therefore S must be +4
- It can easily be written above each example:

  0 0 +4 -2
  - $S(s) + O_2(g) \rightarrow SO_2(g)$
  - Notice how the oxidations state written is the oxidation of ONE element, not taking into account that there are two oxygens.
  - Therefore:
    - Oxidation: $S \rightarrow S^{+4} + 4e^-$
    - Reduction: $2x\ (\frac{1}{2}\ O_2 + 2e^- \rightarrow \frac{1}{2}\ O_2^{-2})$

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A few clarifications….

- In a covalent molecule, if it is unclear which element will take the negative charge, follow electronegativity:
  - For example: with ClF
    - F is more electronegative (is assigned -1)
    - Cl is less electronegative (is assigned +1 to cancel the charge)
    - This is why we write the formula as ClF and not FCl, as Cl behaves as the electron deficient (like a metal)
      - Compounds such as NH_3 may have been established prior to this convention and are therefore backwards (grandfather rule)
- Not everything is ionic!
  - Even though we are assigning oxidation numbers, they are NOT ionic charges so instead of taking or stealing electrons, elements may “gain control” of electrons in a bond.

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Nomenclature in Redox

9.1.3 State the names of compounds using oxidation numbers. (1)

- For compounds that contain elements capable of various oxidation states, the **stock notation** must be used.
  - Roman numerals (I, II, III, IV) are used in this case
    - $FeCl_2$ is Iron (II) chloride
    - $FeCl_3$ is Iron (III) chloride
    - This is used for the transition metals and Pb of group 4
  - This is also used for complex ions
    - $[Fe(CN)_6]^{3-}$ is iron (III) (hexacyanoferrate (III) ion)

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Redox in Practice

9.1.4 Deduce whether an element undergoes oxidation or reduction in reactions using oxidation numbers. (3)

- In order to identify whether a reaction is considered to be a Redox reaction, one can:
  - Deduce all oxidation numbers in the equation (it does not have to be balanced)
  - Examine the see if any of the oxidation number of any elements have changed.
    - An increase in oxidation number (more positive) is Oxidation
    - A decrease in oxidation number (less positive) is Reduction

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Practice Oxidation States

- Determine the oxidation state of each element:

NO_2	$KMnO_4$
N_2O_5	$Fe(OH)_3$
$HClO_3$	$K_2Cr_2O_7$
HNO_3	CO_3^{2-}
$Ca(NO_3)_2$	$K_3Fe(CN)_6$

Using Oxidation Numbers

- An **increase** in the oxidation number indicates that an atom has **lost electrons** and therefore **oxidized.**
- A **decrease** in the oxidation number indicates that an atom has **gained electrons** and therefore **reduced**
- Example

0		+2 +6 -2		+2 +6 -2		0
Zn	+	$CuSO_4$	→	$ZnSO_4$	+	Cu

Zn: **0** → **+ 2** → **Oxidized**

Cu: **+2** → **0** → **Reduced**

Exercise

For each of the following reactions find the element **oxidized** and the element **reduced**

$Cl_2 + KBr \rightarrow KCl + Br_2$

$Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

$HNO_3 + I_2 \rightarrow HIO_3 + NO_2$

- Look on the following slides for answers......

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Exercise 1

For each of the following reactions find the element **oxidized** and the element **reduced**

$$\overset{0}{Cl_2} + \overset{+1\ -1}{KBr} \rightarrow \overset{+1\ -1}{KCl} + \overset{0}{Br_2}$$

Br increases from **–1 to 0 → Oxidized**

Cl decreases from **0 to –1 → Reduced**

K remains unchanged at +1

Exercise 2

For each of the following reactions find the element **oxidized** and the element **reduced**

$$\overset{0}{Cu} + \overset{+1}{H}\overset{+5}{N}\overset{-2}{O_3} \rightarrow \overset{+2}{Cu}(\overset{+5}{N}\overset{-2}{O_3})_2 + \overset{+4}{N}\overset{-2}{O_2} + \overset{+1}{H_2}\overset{-2}{O}$$

Cu increases from 0 to +2 → **Oxidized**

Some N in nitric acid from +5 to +4 → **Reduced**

The nitrogen that ends up in copper nitrate remains unchanged, same for hydrogen and oxygen

Exercise 3

For each of the following reactions find the element **oxidized** and the element **reduced**

$$\overset{1\ \mathbf{+5}\ -2}{HNO_3} + \overset{\mathbf{0}}{I_2} \rightarrow \overset{+1\ \mathbf{+5}\ -2}{HIO_3} + \overset{\mathbf{+4}\ -2}{NO_2}$$

N is reduced from +5 to +4 → **Reduced**

I is increased from 0 to +5 → **Oxidized**

The hydrogen and oxygen remain unchanged.

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Half-Equation Examples:

$$\overset{0}{Cl_2} + \overset{+1\ -1}{KBr} \rightarrow \overset{+1\ -1}{KCl} + \overset{0}{Br_2}$$

Oxidation ½ Rxn: $Br^{-1} \rightarrow Br + 1e^-$

Reduction ½ Rxn: $Cl + 1e^- \rightarrow Cl^{-1}$

$$\overset{0}{Cu} + \overset{+1\ +5\ -2}{HNO_3} \rightarrow \overset{+2\ \ +5\ \ -2}{Cu(NO_3)_2} + \overset{+4\ -2}{NO_2} + \overset{+1\ \ -2}{H_2O}$$

Oxidation ½ Rxn: $Cu \rightarrow Cu^{2+} + 2e^-$

Reduction ½ Rxn: $N^{+5} + 1e^- \rightarrow N^{+4}$

$$\overset{1\ +5\ -2}{HNO_3} + \overset{0}{I_2} \rightarrow \overset{+1\ +5\ -2}{HIO_3} + \overset{+4\ -2}{NO_2}$$

- Oxidation ½ Rxn: $I \rightarrow I^{+5} + 5e^-$
- Reduction ½ Rxn: $N^{+5} + 1e^- \rightarrow N^{+4}$

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Oxidation-Reduction Reactions

- All redox reactions have at least one element **oxidized** and at least one element **reduced**
- Occasionally the same element may undergo both oxidation and reduction.
  - This is known as an **auto-oxidation reduction**
  - IB refers to it as **Disproportionation**
  - A common example is the reaction between Cl_2 and H_2O

$$\overset{0}{Cl_2} + \overset{+1}{H_2}\overset{-2}{O} \rightarrow \overset{+1}{H}\overset{-2}{O}\overset{+1}{Cl} + \overset{+1}{H}\overset{-1}{Cl}$$

- $Cl_2 + H_2O \rightarrow HOCl + HCl$
  - Oxidation: $Cl \rightarrow Cl^{+1} + 1e^-$
  - Reduction: $Cl + 1e^- \rightarrow Cl^{-1}$